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STRATEGIC PETROLEUM RESERVE

An Analysis of Potential Corrosion Problems

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SUMMARY

Discussions were held with personnel associated with the **SPR** Program in order to document the corrosion problems observed to date as well as to determine the conditions presumed to exist within the system components. An **analysis** was then made using first principles to ascertain the severity or likelihood of corrosion. Four major system components were identified and considered in detail. Conclusions relating to each system are as follows:

A. Interior Pipeline Corrosion by Brine

Concentrated brines are less corrosive than dilute brines. Corrosion rates are acceptably **low** if the brines are oxygen free. Numerous drill strings have lasted over twenty years in concentrated brines, but perforations have occurred during leaching operations (German Program) when dilute brines containing oxygen were involved. The major potential problem will involve corrosion during leaching of new caverns.

B. Exterior **Pipeline** Corrosion by Surface Soil Water

This aspect of corrosion has been successfully handled by the petroleum industry for many years by cathodic protection. Present **SPR** specifications call for cathodic protection, and there should be no problems if accepted practices are followed.

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C. Exterior Downhole Casing Corrosion

Acidic groundwaters containing dissolved sulfur species represent a potentially severe corrosion problem. The problem will be most severe at sites previously used for sulfur recovery. General (uniform) corrosion rates may be as much as three times greater when dissolved H_2S is present compared to dissolved oxygen. The worst aspect is the possibility of pitting corrosion (highly localized) which has probably caused casings to perforate within two years. Cathodic protection is difficult to utilize, and the best way to circumvent pitting corrosion may be to use an alloy steel such as a ferritic stainless. However, this is costly and would require laboratory tests to establish if the required protection can be achieved for the existing conditions.

D. Galvanic Corrosion

Dissimilar metals electrically coupled (in intimate contact with each other) will form a battery, and the most active will corrode preferentially, especially in brine which is a good electrolyte. Present specifications require isolation flanges which would act as dielectrics. The use of dielectric couplings between all dissimilar metals or the use of similar metals throughout will effectively eliminate galvanic corrosion.

Corrosion has not been a major problem thus far, but there are two areas in which corrosion could be critical: corrosion of casings during dome leaching, and exterior corrosion of well casings by acidic groundwaters.

INTRODUCTION

One of the prime technical considerations for the viability and reliability of the SPR program involves the compatibility of the system components with their environment. In this case the concern is aqueous corrosion. Part of the charter of the Sandia program to provide technical assistance to the SPR concerns an analysis of corrosion problems that either exist presently or might arise in the future and to address their severity with recommendations to alleviate or to minimize corrosion. This task has been undertaken by a two-fold approach. The first endeavor utilized extensive discussions with various personnel from DOE in New Orleans, **Parsons-Gilbane Co.**, **Parsons-Brinkerhoff Co.**, and Dow Chemical Co. to determine from their experience what problems exist. The second endeavor was to use applicable data from the literature and to apply first principles of corrosion to assess the magnitude of the problems. The following report summarizes what was learned and what might be expected. It became apparent from the discussions that the corrosion analysis should be considered in terms of the following system components: a) interior pipeline corrosion by brine, b) exterior pipeline corrosion by surface soil water, c) exterior **downhole** casing corrosion by acid waters in the **caprock**, and d) galvanic corrosion of equipment. Each of these cases will be considered individually.

INTERIOR PIPELINE CORROSION BY BRINE

Results of Discussions

The general concensus is that corrosion by brine does not represent a problem as long as oxygen is kept out of the brine. The oxygen content of the brines displaced from the caverns has been measured and is generally

1 ppm or less. As long as air is not introduced into the pipes or the brine aerated in any fashion, corrosion is minimal (N. Shourbaj, **Parsons-Gilbane**; H. **Gomm**, Parsons-Brinkerhoff; Mayeaux, DOE; B. Henderson, Dow Chemical; and R. Young, Dow Chemical). Mr. Bailey Henderson, a **Dow** Chemical Co. representative at Bryan Mound, reports that plain-carbon steel strings in the caverns have survived 25 years use during which concentrated brine was pumped out by Dow Chemical Co.

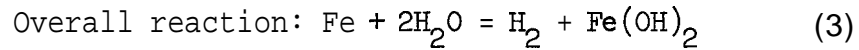
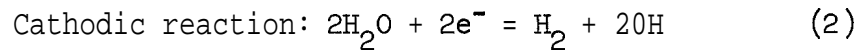
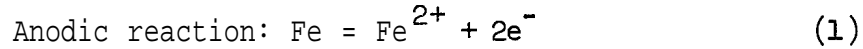
The only case of excessive corrosion in brine was reported by H. Gomm which concerned dilute brines during dome leaching in the German storage program. As will be shown subsequently, dilute brines are much more corrosive than concentrated brines, and the oxygen content of the brine (perhaps the most important factor) was much higher because seawater was pumped into the domes for leaching. Perforation of the casings occurred in about two years, and it was necessary to replace the casings.

Mr. Robert Young, **Dow Chemical** Co. presently at Stratton Ridge, reports that corrosion test coupons were installed in pipes used for transporting concentrated brines after removal from the underground caverns. The test coupons were utilized for several years and were removed periodically for weighing. Mr. Young was involved in the test-coupon program but states that no records were kept. Corrosion rates of 0.005 to 0.006" per year were measured. Mr. Vincent Lepardo, project manager for **Parsons-Brinkerhoff**, stated that a corrosion allowance of 0.05 to 0.1" was designed into the pipeline for its anticipated lifetime of 30 years.

Mr. Young also reported that the only documented case of excessive corrosion in brine pipelines has occurred when the brine becomes supersaturated and precipitates out salt crystals. Corrosion was observed to occur in such cases under the crystal deposits due to a concentration cell effect.

Fundamental Analysis

Steel corrodes in water according to the following reaction when oxygen is not present:



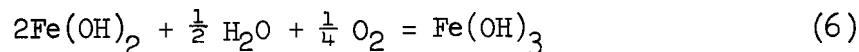
Reaction (2) is slow, and thus the overall reaction rate is slow. However, if oxygen exists, an alternative cathodic reaction may occur which is much more rapid than reaction (2):



which combined with reaction (1) gives the following overall reaction:



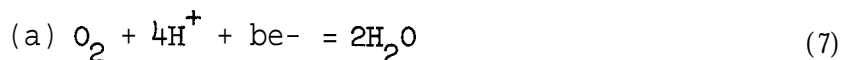
The ferrous hydroxide is further oxidized to form rust by the reaction



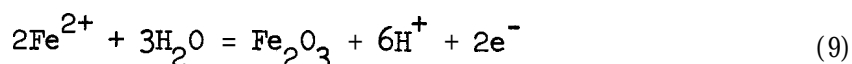
Reactions (5) and (6) are much faster than reaction (3), and steel will corrode much more rapidly in brines containing oxygen than in oxygen-free brines.

The corrosion of steel depends upon the pH, temperature, Cl^- content as well as on the oxygen content of the solution. In order to understand this behavior, it is useful to utilize a **Pourbaix** diagram, a thermodynamic stability diagram which is a plot of potential versus pH. Each combination of these variables leads to the formation of various compounds or species which can be mapped to show whether the metal is immune from corrosion, whether it corrodes, or whether a reaction product film forms which may passivate the metal. Figure 1 shows such a diagram for iron and water in the absence of chlorides at 25°C (steel is essentially iron with some minor additions). If the pH potential combination falls within the iron region (Fe), the metal is immune from corrosion. If the combination falls within the Fe_3O_4 or Fe_2O_3

fields,* the iron may become passivated if the reaction product film is impervious and does not continuously dissolve or spall. Lines (a) and (b) represent the stability limits for water, the respective reactions being



It should also be noted that the lines for the stability regions depend upon the concentration of dissolved species as shown by the parallel lines for reactions (23) and (28). For **example**, reaction (28)



has a potential dependence given by

$$E_o = 0.728 - 0.1773 \text{ pH} - 0.0591 \log (\text{Fe}^{2+}). \quad (10)$$

A low concentration of Fe^{2+} expands the "corrosion" region markedly. The equilibrium potential of $\text{Fe}-\text{Fe}^{2+}$ is -0.62 volt, which for a neutral solution and low Fe^{2+} content places iron in the "corrosion" region. The **Pourbaix** diagram will be used to show how **pH** and chloride content change the nature of the corrosion products.

Corrosion is an electrochemical process, and the current is therefore a measure of the corrosion rate. If steel is anodically polarized (held at a given potential during which time the current is measured), the corrosion rate increases with increasing potential until the potential of the passivity region is reached. The current then drops markedly and corrosion is suppressed. The iron remains passivated until much higher potentials are reached at which point the passive film becomes unstable and breaks down locally (the **trans-passive** region). Corrosion then proceeds with further increases of potential.

* It has been shown that the passive films formed on iron are the oxides. ^{2,3} However, the **Pourbaix** diagrams can also be drawn to show the equilibria involving the iron hydroxides instead of the oxides.

This phenomenon is illustrated in Fig. 2 for five different pH values.⁴ It may be noted that the passivity region increases with increasing pH. The pH of the brines in the SPR program vary from about 6.2 to 7.9 with the exception of Bryan Mound Well No. 4 which has a value of 10.4.

Dilute salt solutions have a highly deleterious effect on corrosion by markedly reducing the potential at which transpassivity occurs. This effect is shown in Fig. 3⁵ for the same pH values depicted in Fig. 2. The presence of chlorides enables corrosion to occur over a much wider range of conditions than in chloride-free solutions. In other words, the protective nature of product films is lost or greatly reduced in the presence of chlorides.

Reaction (4), oxygen depolarization, controls the corrosion rate in neutral chloride solutions, and thus with the chlorides rendering the passivation films less protective, one expects that the corrosion rate should increase with chloride content. This does indeed occur up to about 3 wt.% NaCl (seawater concentration), but at high concentrations the rate decreases to values considerably less than for chloride-free water containing oxygen⁶ as shown in Fig. 4. The decrease results from a significant decrease in oxygen solubility with increasing chloride content.

According to Uhlig⁶ the increased corrosion rate observed with increasing chloride level in the dilute range results from the increased conductivity of the chloride solution. High-purity water has a low conductivity, and thus anodes and cathodes must be near one another. Hydroxyl ions forming at the cathodes are then always near Fe^{+2} ions forming at the anodes, the hydroxyl and iron ions reacting adjacent to the surface to form an adherent layer of rust which is moderately protective. The high conductivity chloride solutions may have anodes and cathodes which are widely spaced. The ions formed at each electrode

can diffuse away **and** react with each other to form flocculent rust away from the metal surface. This type of rust does not offer any protection, and the dissolved oxygen has easy access to the cathodic areas on the iron surface to cause additional corrosion.

The **Pourbaix** diagrams represent thermodynamic stabilities and offer nothing concerning reaction kinetics. A compilation of rate **data**⁸ have been superimposed on the **Pourbaix** diagram for iron and are shown in Fig. 5 for oxygen-free, stirred solutions. It is clear that low **pH** and high potential increase the corrosion rate. Fortunately, the **pH** potential conditions of brine places iron in a region of low corrosion rate, which accounts for the observed long lifetime of steel strings immersed in the brine caverns.

Conclusions

The stated facts: 1) concentrated brines are less corrosive than dilute ones, 2) the corrosion rates are low for all brines as long as the brines are oxygen-free, and 3) some strings in the brine caverns have lasted for more **than** 20 years, agree with **all** aspects of corrosion theory. Corrosion has not been a problem in the past and should not be one in the future as long as air is kept out of the system. The latter is by far the most important aspect of brine-pipeline compatibility.

EXTERIOR PIPELINE CORROSION BY SURFACE SOIL **WATER**

Results of Discussions

Buried pipelines have been used satisfactorily by the petroleum industry for a number of years by cathodically protecting the pipes. This practice is being used on the **SFR** program. Mr. Vincent Lepardo, Project Manager, **Parsons-Brinkerhoff**, stated that the ANSI **B31-4** specification (liquid petroleum piping code) is being followed. Mr. Murray Wiggins, DOE, reported that some pipes are coated with an epoxy and cathodically protected. However, the

protection is not necessarily applied **immediately**. Test leads from the pipes at various intervals protrude from the soil and are used to measure the emf against the soil. After settling and compaction of the soil cathodic protection is applied if the emf measurements so indicate. The emf, of course, depends on the nature of the local soil. The time period from installation to protection, if needed, is about 1 year. The pipes generally have a wall thickness of about 0.625 inch, and the tolerable corrosion either with or without cathodic protection is 0.02"/year.

Mr. Paul Mayeaux, Assistant Site Manager, Bryan Mound, stated that a 30" pipe from the seaway dock to the Jones Creek Tank Farm (about 7 miles long), has been in place for about one year and has not been protected, although it is planned to use impressed currents for cathodic protection. This pipe is adjacent to a 20" pipeline (about 10' separation), and a single rectifying system for the two pipes will be used. Mayeaux further states that the DOT regulation, Docket HM6, Title 49, Chapter 6, paragraph 195.414, sub-paragraph (f) operation and paragraph 195.242, sub-paragraph Construction, will be followed. It should be noted that the 30" pipeline will be transporting supersaturated brine from the settling pond at ambient temperature. This brine will undoubtedly be aerated, but the oxygen **solubility** in a brine containing 25-26% NaCl should be quite low,⁷ and thus internal corrosion should not be severe.

Fundamental Analysis

Cathodic protection is by far the best means of protecting structural members in corrosive environments, e.g., ships in seawater, pipelines, etc. There is a long history for this technology, and it works well. Referring to the **Pourbaix** diagram once again, one observes that if the potential of

iron in the particular electrolyte can be shifted out of the corrosion region, protection can be achieved. The best means for doing this is shown in Fig. 6 which illustrates three possibilities including cathodic protection. Either sacrificial anodes (zinc, aluminum, magnesium) or impressed direct current can be used to cathodically protect the pipe.

Conclusions

External surface corrosion of buried pipelines can be completely suppressed or reduced to tolerably low levels by cathodic protection. This technique has been used successfully for many years by the petroleum industry and is being utilized on the SPR program. If suitable practices are followed, corrosion of buried pipelines should not be a problem.

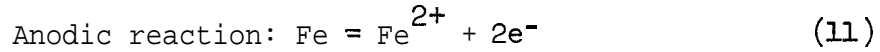
EXTERIOR ~~DOWNHOLE~~ CASING CORROSION

Results of Discussion

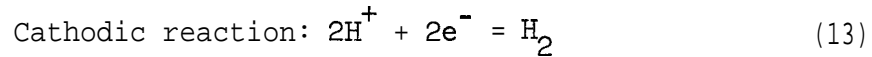
The exterior corrosion of casings can be significant according to Mr. Robert Young, Dow Chemical Co. at Stratton Ridge. The problem arises whenever the **caprock** contains a sulfur species and is **almost** always associated with sulfur mining activities. The casings perforate in extreme cases and may do so in about 2 years, depending upon the soil-rock severity. It has been Dow practice to simply insert a new pipe as a liner inside the perforated casing. Obviously, there is a limit to the number of times this can be done. It is not possible to cathodically protect the casings because they are cemented in place. It was further stated that welded pipe exhibits preferential corrosion in the heat affected weld zone, and thus **seamless** pipe is preferred.

Fundamental Analysis

Exterior casing corrosion by waters appears to be a sulfide corrosion problem involving dissolved H_2S . The electrode reactions in the corrosion process and their respective potentials are given by:⁹



$$\text{Anode Potential: } E_A = -0.440 + 0.0592 \log a_{Fe^{2+}} \quad (12)$$



$$\begin{aligned} \text{Cathodic potential: } E_C = & 0.0592 \log [H^+] \\ & - \frac{0.0592}{2} \log P_{H_2} \end{aligned} \quad (14)$$

The driving force for the corrosion reaction is the difference between reactions (13) and (11)

$$E_C - E_A = 0.440 + \frac{0.0592}{2} \log \frac{[H^+]^2}{[Fe^{2+}] P_{H_2}} \quad (15)$$

The reaction potential depends on the activity of hydrogen and iron ions in solution and the pressure of gaseous hydrogen. The ferrous ion content is a function of the solubility product of ferrous sulfide which is given by¹⁰

$$[Fe^{2+}] [S^{2-}] = 10^{-19} \text{ at } 25^\circ C \quad (16)$$

The sulfide ion content depends on the dissociation constants of hydrogen sulfide

$$\frac{[H^+] [HS^-]}{[H_2S]} = 1.15 \times 10^{-7} \quad (17)$$

$$\frac{[H^+] [S^{2-}]}{[HS^-]} = 10^{-15} \quad (18)$$

Multiplication of equations (17) and (18) gives

$$\frac{[H^+]^2 [S^{2-}]}{[H_2S]} = 1.15 \times 10^{-22} \quad (19)$$

Substituting equations (16) and (19) into equation (15) gives an expression based on thermodynamics for the corrosion potential* when H_2S exists in the water

$$E_C - E_A = 0.353 + 0.0296 \log \frac{[H_2S]}{P_{H_2}} \quad (20)$$

The driving potential thus depends only on the hydrogen sulfide activity and on the hydrogen partial pressure. The corrosion rate increases with increasing hydrogen sulfide content as seen in Fig. 7,¹¹ which also shows the effect of increasing oxygen and carbon dioxide contents. Note the different horizontal scales for each dissolved gas. Taking the maximum values for H_2S and the solubility limit of oxygen (about 8 ppm), one can observe that the corrosion rate due to H_2S may be more than three times greater than that due to oxygen, which had previously been shown to be highly deleterious.

Although reaction (20) indicates that the corrosion should be independent of pH, it is not the case as seen in Mg. 8.¹² Corrosion is less at higher pH values. A possible explanation for this effect is as follows. The activities of sulfide and iron ions in a solution saturated with ferrous sulfide and $P_{H_2S} = 1$ atm are given by:

$$\begin{aligned} [S^{2-}] &= \frac{1.17 \times 10^{-23}}{[H^+]^2} \\ [Fe^{2+}] &= \frac{[H^+]^2}{1.17 \times 10^{-4}} \end{aligned} \quad \begin{array}{l} w \\ (22) \end{array}$$

Thus, the ferrous ions are in great excess at low pH values, whereas, at high pH values the sulfide ions are in excess. During corrosion ferrous ions form at the anodes and diffuse into the solution and sulfide ions diffuse towards the anode. Precipitation of ferrous sulfide forms when the solubility

*In some cases kinetics of the reaction may determine the actual corrosion potential.

product is exceeded. At low pH the ferrous ions far exceed the sulfide ions, and the few sulfide ions near the anode precipitate out to form ferrous sulfide: and the precipitation zone then moves out into the solution. The net result is that the precipitate does not form a protective film. On the other hand at high pH the sulfide ions far exceed the ferrous ions, and any ions forming at the anode precipitate out immediately with the available sulfide ions adjacent to the surface. A more protective film results in high pH solutions.

All previous comments concerning corrosion of steel in solutions containing H_2S applied to general (uniform) corrosion. If the corrosion rates for uniform corrosion are well-known, it is possible to design systems with a given lifetime on the basis that a known amount of metal will be consumed. On the other hand, localized corrosion occurs in the presence of H_2S in the form of either pitting or stress corrosion cracking (SCC). The localized corrosion is an insidious phenomenon, and it is nearly impossible to predict rates and component lifetimes. Although it is not known, intuitively it would appear that the casing perforations observed by Dow Chemical Co. over the years occurred by pitting corrosion rather than by a uniform attack. Although SCC is well documented throughout the petroleum industry, there is no evidence that SCC has occurred in the chemical industry where brine removal from sub-surface domes has occurred.

There are some reasons why SCC has not been observed in brine systems and may not be observed in the SPR program. First, a stress is required which is an appreciable fraction of the yield strength. Generally, the higher the stress the shorter the time for failure. The susceptibility increases with the strength of material. SCC usually occurs in deep sour wells which

*The sulfide concentration drops adjacent to the precipitate. The low-concentration zone thickness depends on the amount of convection and/or stirring in the electrolyte.

exhibit high H_2S levels and which are highly stressed due to the depth of the wells. This is not the case for SPR. The H_2S levels are much lower, and the shallow wells are not highly stressed. Second, the susceptibility to SCC increases with the strength of the steel. The low-carbon steels used in SPR are the weakest steels available and have a low inherent susceptibility. These steels **can** be made to crack under certain conditions, but these conditions do not exist in the SPR program.

Pitting of steel in H_2S solutions has been studied by Dvoracek¹³ over a range of **pH** and potential combinations and for various surface preparations. Slightly oxidized steel, simulating mill scale formed during the processing of pipe, was found to pit in solutions containing 355 ppm of **NaCl** and saturated with H_2S . The pitting regime is shown on a **pH potential** diagram in Fig. 9. The pitting potential is independent of **pH** between 3 and 6. At higher values of **pH** the pitting potential becomes more noble, and pitting does not occur as readily as at low **pH** values. The line delineating the "**pitting-no pitting**" transition may be explained on the basis of the species present as noted in Figure 10.¹⁴ The primary species between **pH** of 3 to 6 is H_2S , whereas at higher **pH** the primary species is HS^- . It should be noted that pitting requires the presence of chloride ions. Thus, any spillage of brine around well heads would cause seepage into the ground and would eventually lead to the presence of chloride in the groundwaters. It appears that those salt domes which had been mined previously for sulfur would be sites at which pitting of well casings into the dome might occur.

Conclusions

The corrosion of casing exteriors by acidic groundwaters is the most severe corrosion problem in the SPR program and may be attributed to the presence of dissolved sulfur species. The problem will be particularly bad

at sites containing large amounts of sulfur. The worst possible case predicts that the corrosion rate by dissolved H_2S can be three times greater than that due to dissolved oxygen (the maximum rates occur at different concentrations levels for the two gases). The corrosion rate increases markedly with decreasing pH.

Past experience shows that in extreme cases well casings may perforate in approximately two years. It has been customary to insert new casings inside the perforated ones. However, this technique is obviously not satisfactory, and the reduced diameter of new inserts would reduce the flow rates at which oil could be removed.

Stress corrosion cracking should not be a problem, but pitting corrosion could be a major problem. It is thought that the observed perforations by Dow Chemical Co. are due to pitting in groundwaters containing dissolved H_2S (or some other sulfur species) and chloride.

Cathodic protection is not a very practical method of minimizing corrosion, although it might conceivably be employed. A much more reliable cure, but one which would be very expensive, would be to use a steel which does not pit under the existing conditions, i.e., a ferritic stainless steel.

GALVANIC CORROSION

Results of Discussions

Although good corrosion engineering practice avoids galvanic cells in systems exposed to corrosive liquids, galvanic cells do occur occasionally and must be guarded against. An instance which did occur in the German program was mentioned by H. Gomm, Parsons-Brinkerhoff. A stainless steel valve in a cast iron seat was used in a brine line, and the seat underwent

gross galvanic corrosion. Although this was the only reported case, others can be expected whenever dissimilar metals are used without being electrically insulated from one another.

The possibility of galvanic corrosion has been recognized as evidenced by the pipeline specifications which call for isolation flanges according to V. Lepardo, Parsons-Brinkerhoff, and P. Mayeaux, Asst. Site Manager, Bryan Mound.

Fundamental Analysis

Each metal undergoes an electrode reaction during corrosion which gives rise to a potential as noted in Equation (12) for iron. The voltage is measured against a standard hydrogen electrode (0.00 volt) and will depend upon the electrolyte to a limited degree. Some potentials measured in neutral seawater are summarized in Table I.

TABLE I
Galvanic Series for Some Metals and Alloys in
Air-Saturated Neutral Seawater

<u>Metal</u>	<u>Potential (volt)</u>
Magnesium	-1.32
Zinc	-0.78
Aluminum	-0.67
Mild steel	-0.40
Cast iron	-0.35
13% Cr steel (active)	-0.30 approx.
304 stainless (active)	-0.30 approx.
Lead	-0.26
Brass (60 Cu-40 Zn)	-0.07
Copper	+0.10
Cupronickel (70 Cu-30 Ni)	+0.34
13% Cr steel (passive)	+0.40 approx.
304 stainless (passive)	+0.40 approx.

If any two metals having different electrode potentials are electrically coupled to each other, the one lying above the other in the galvanic series is the anode and will corrode preferentially. The anodic metal will cathodically protect the metal which is lower in the series. The greater the potential difference between the two metals, the greater the driving force for corrosion. The top of the table represents the "active" metals, whereas the bottom represents the "passive" or noble metals. Note that stainless steels can shift their positions in the series markedly by **passivation**. If iron and either **13%** Cr steel (ferritic stainless) or **304** stainless are coupled, and the stainless is not passivated, a small potential difference results (approx. 0.10 volt), and corrosion of the iron will be minimal. On the other hand, if the stainless becomes passive, the potential difference is increased to about 0.80 volt, and corrosion of iron is accelerated.

There are two other factors in addition to the potential difference that determine the extent of galvanic corrosion. These are the relative areas of the anode and cathode and the conductivity of the solution. If the anodic area is small compared to the cathodic area, the current density at the anode becomes very high, and the local corrosion rate of the anode **may** be excessive. On the other hand, if the cathodic area is small relative to that of the anode, the corrosion rate will be much lower. As far as conductivity is concerned, brine is an excellent electrolyte, and charge transport is no problem. In other words, if a galvanic couple exists in brine, the anodic metal will corrode.

Conclusions

Galvanic corrosion can be expected whenever dissimilar metals are electrically coupled (in intimate contact with each other) in brine of all concentrations. Corrosion will be greater if large differences exist between

the electrode potentials of the coupled metals and if the anodic area is small compared to the cathodic area. This form of corrosion can be minimized by using the same metals for various components or by electrically insulating the dissimilar metals.

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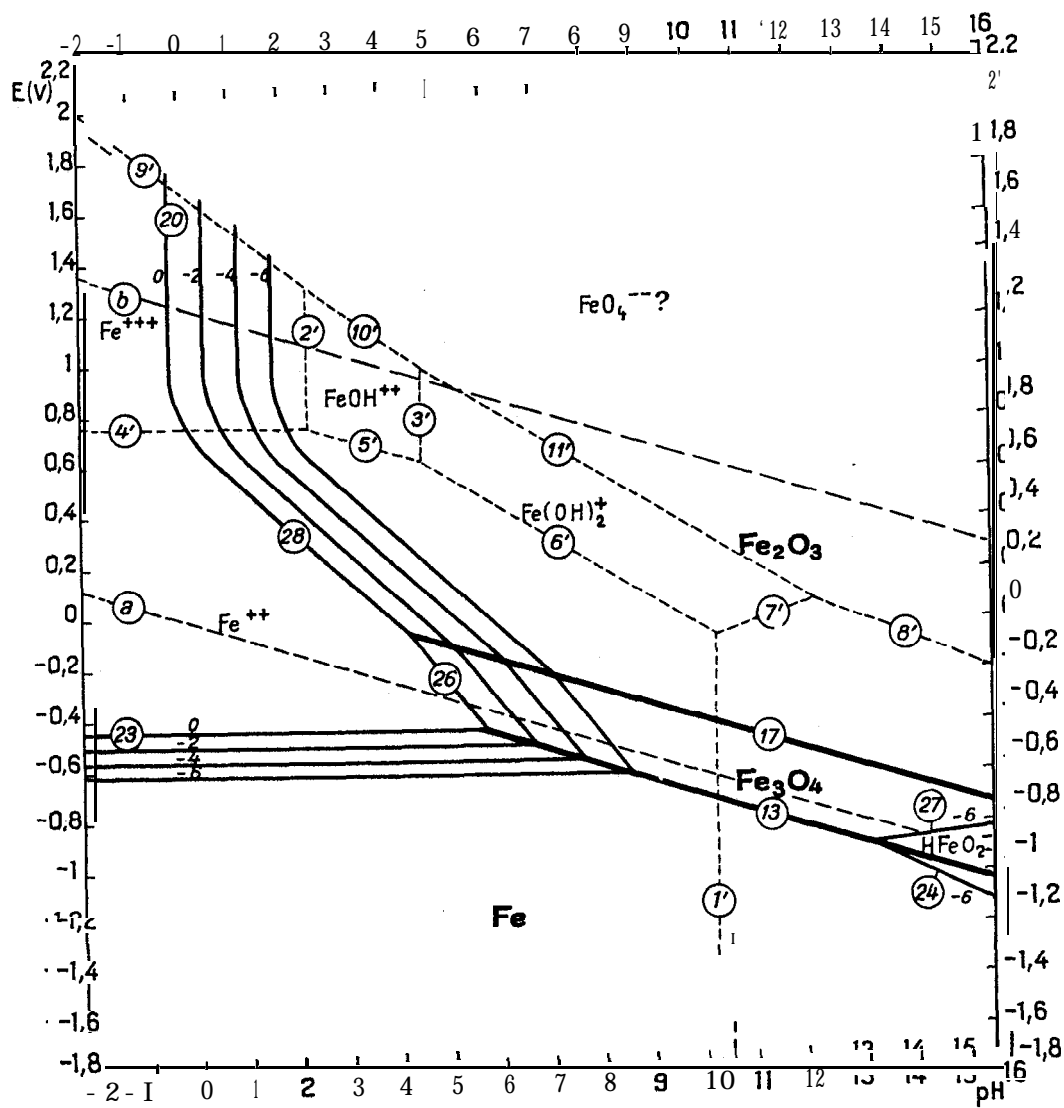


Figure 1. **Potential-pH** Equilibrium Diagram for the System Iron-Water at 25°C.

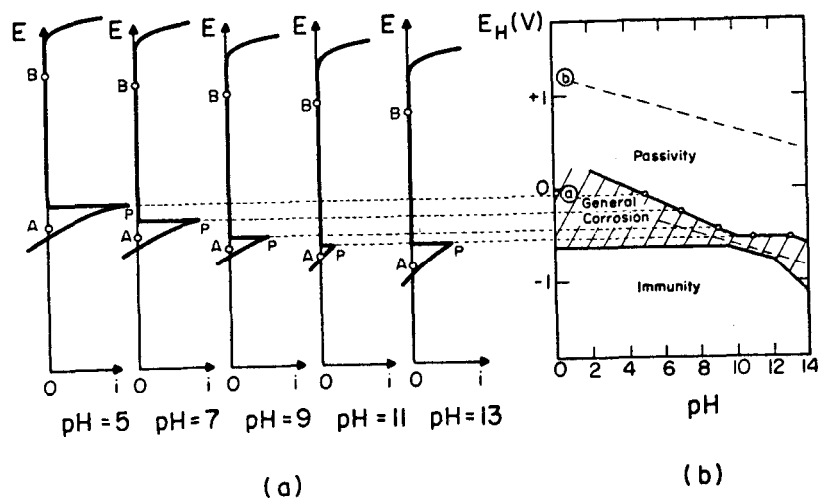


Figure 2. Behavior of Iron in Chloride-Free Solutions: (a) **Polarization** Curves in Solutions of Various pH; (b) Experimental Conditions of Corrosion, Immunity and Passivity.

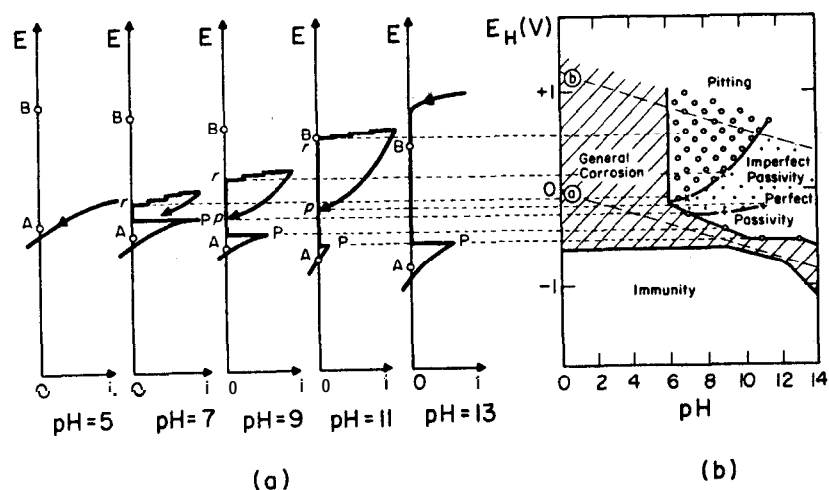


Figure 3. Behavior of Iron in Solutions Containing 355 ppm NaCl: (a) **Polarization** Curves for Solutions of Various pH (same values as in Fig. 2; (b) Experimental Conditions of Corrosion, Immunity, Perfect and Imperfect Passivity, and Pitting.

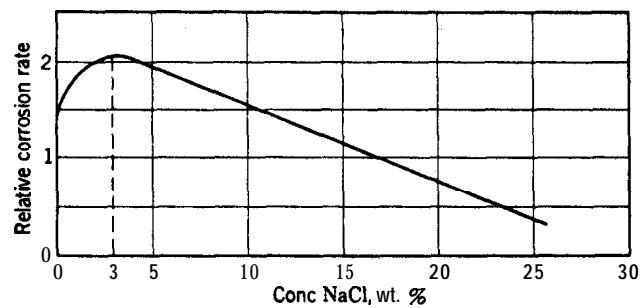


Figure 4. Effect of Sodium Chloride Content on the Corrosion of Iron in Aerated Solutions at Room Temperature.

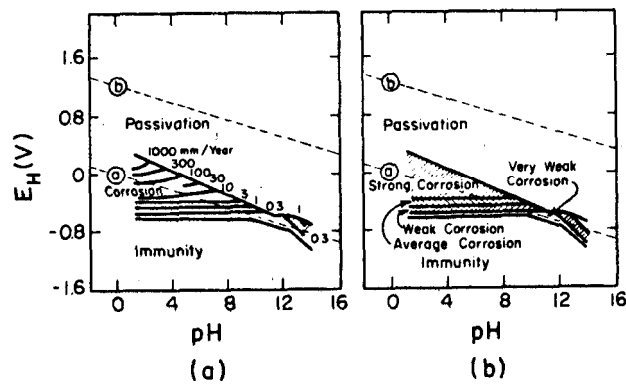


Figure 5. Experimental. Conditions of Corrosion, Immunity, and **Passivity** of Iron in Stirred Aqueous Solutions: (a) Corrosion Rates; (b) Corrosion Intensity.

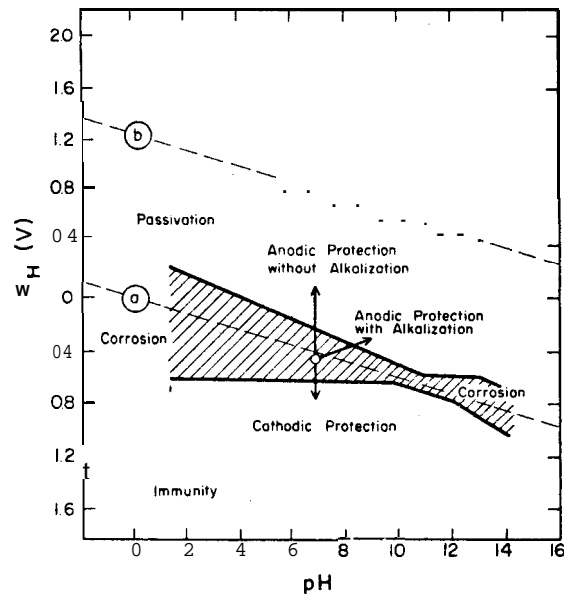


Figure 6. Various Means of Corrosion Protection for Iron.

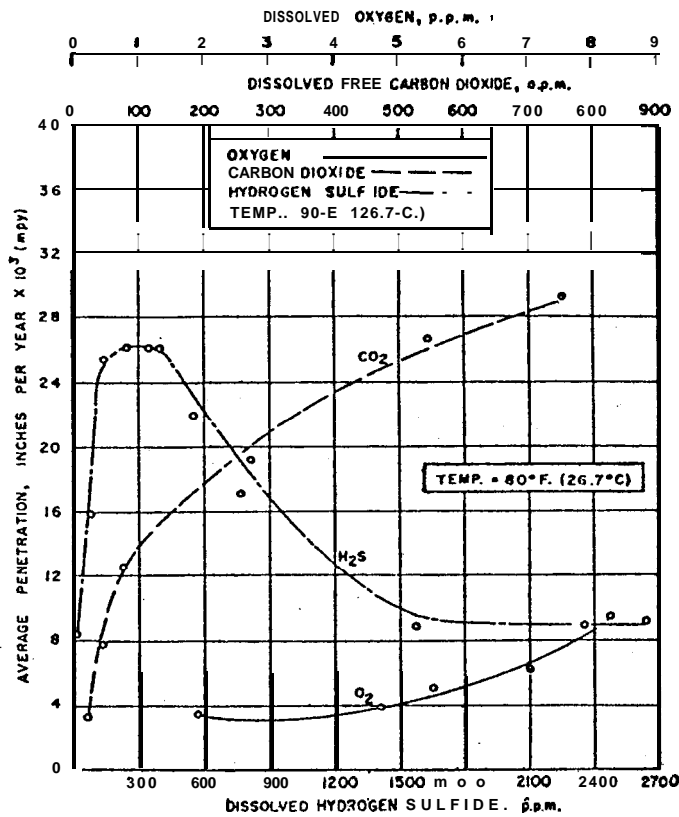


Figure 7. Effect of Dissolved Gases on Steel Corrosion. Note Different Concentration Scales for Each Gas.

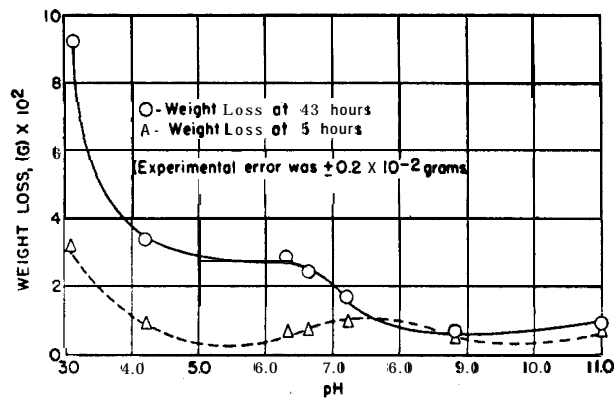


Figure 8. Weight Loss of Iron Versus pH in an Aqueous Solution Containing H_2S .

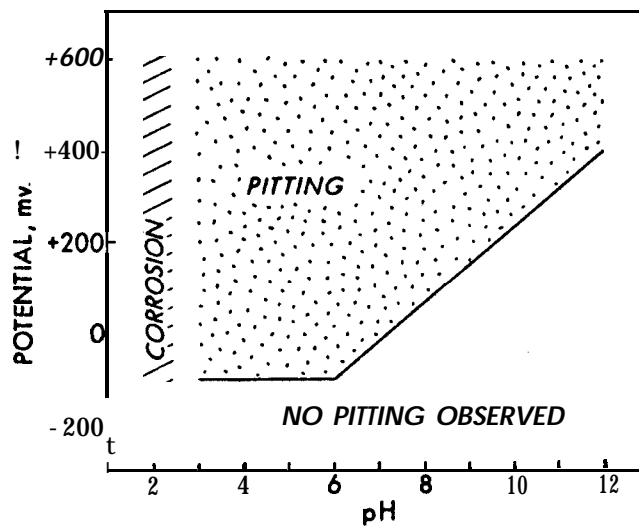


Figure 9. Potential pH Diagram of Oxide-Coated Steel in Water Containing 355 ppm NaCl and Saturated with H_2S .

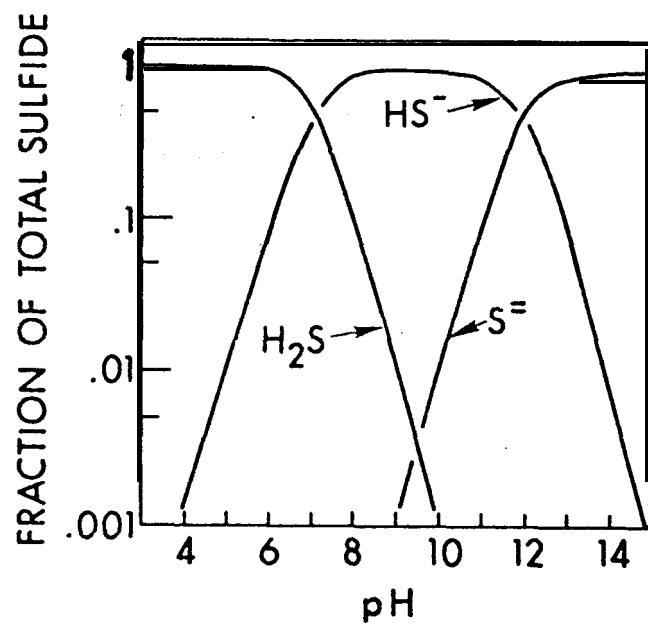


Figure 10. Distribution of Various Species in Water Containing Dissolved H_2S .

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